Atomic Structure

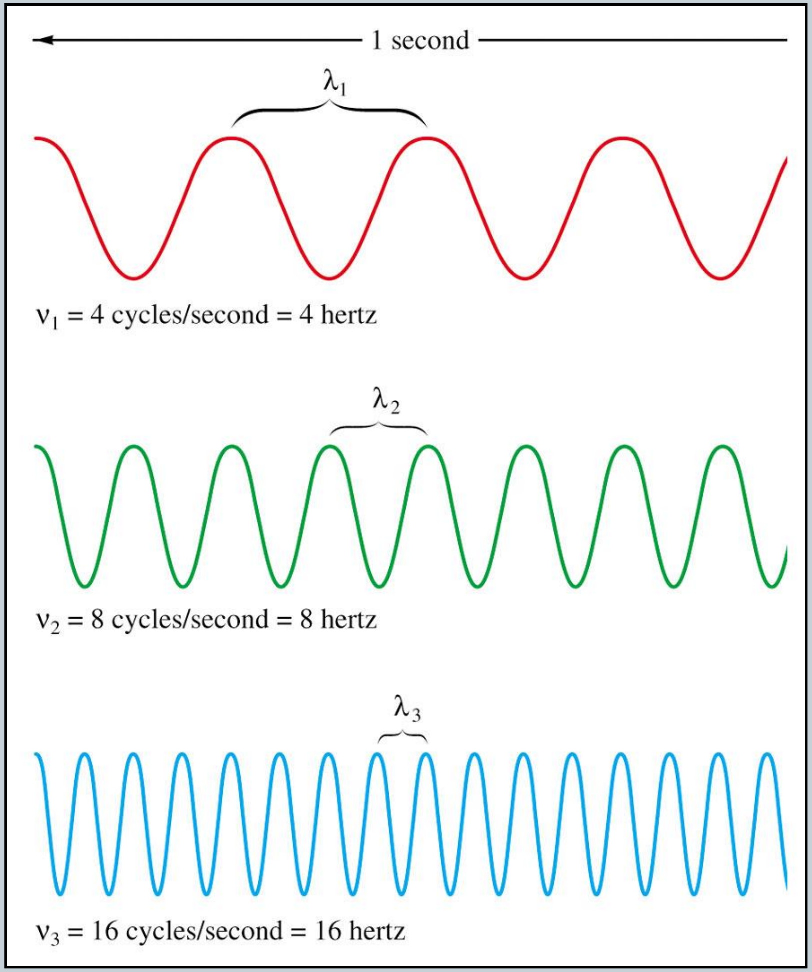

J.J. Thomson
Robert Millikan
Ernest Rutherford
Atomic Structure

Electromagnetic Radiation

- Its interaction with matter—chemists use this to study the structure of the atom
- Consists of waves, which have three characteristics:
  - Wavelength (λ)—distance between two consecutive peaks or troughs of a wave
  - Frequency (f)—number of waves (cycles) per second that pass a given point in space
    - Measured in sec^{-1} (Hz)
  - Speed (c)—3.00×10^8 m/s in a vacuum
- All related through the expression:

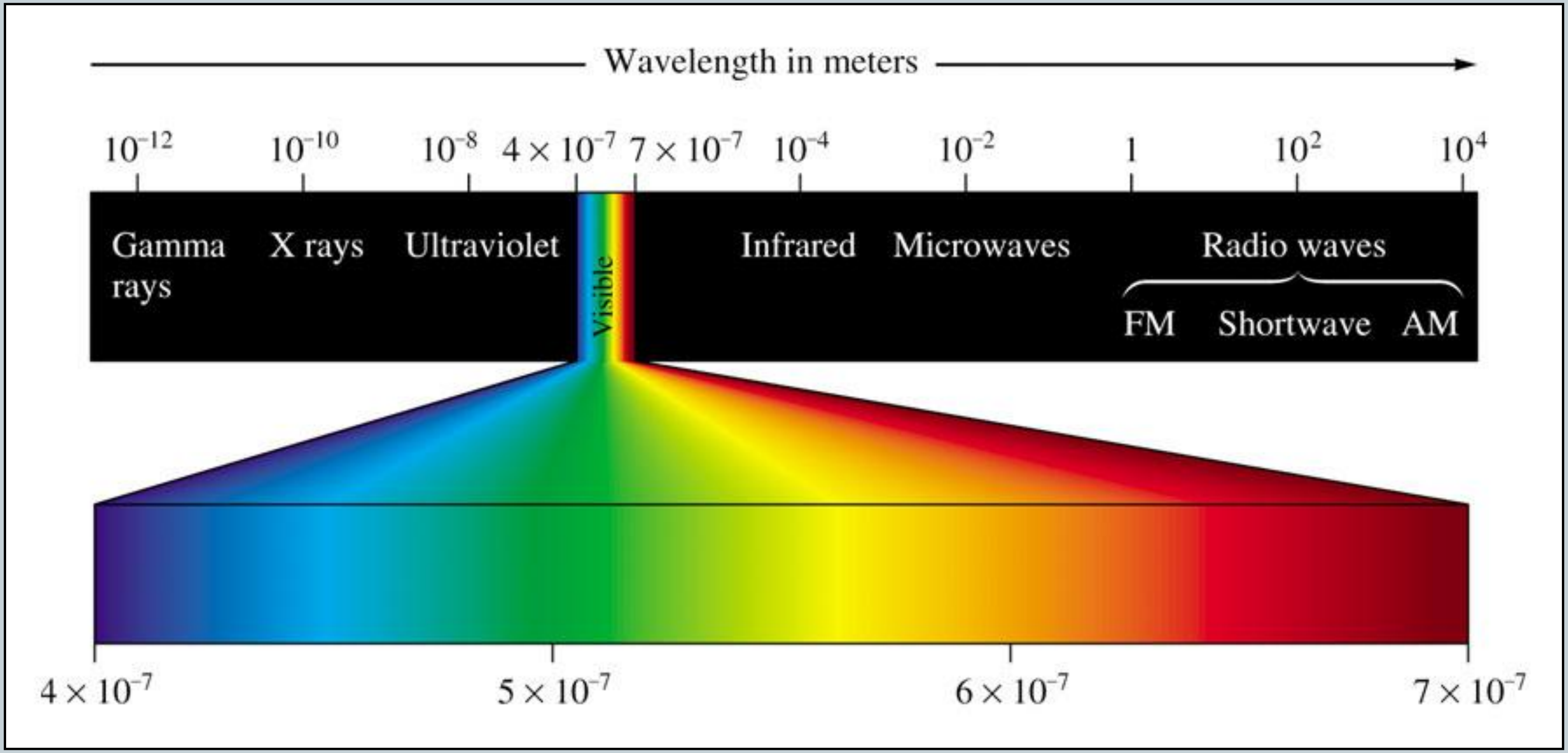
$$\lambda f = c$$

Characteristics of Waves



Electromagnetic Radiation

- Classified into seven categories based on wavelength



Prior to 1900

Energy (waves) ≠ Matter (particles)

Max Plank & His Contributions

- Proposed that energy is emitted and absorbed in packets called quanta
  - $E_{proton} = hf$ where $h = 6.626 \times 10^{-34}$ J·s
  - Energy is proportional to frequency
- Conclusion: energy has particulate properties

Albert Einstein & His Contributions

- Called the particles photons
- Based on Planck's relationship and relationship between wavelength and frequency, he proposed the following relationship:
  - $E_{proton} = hf = \frac{hc}{\lambda}$
  - Proposed that all electromagnetic radiation is quantized
- Photoelectric effect—electrons are emitted from the surface of a metal when light strikes it above a particular frequency
  - Proof that energy is quantized
  - Energy also has mass ($E = mc^2$)

Dual Nature of Light

- Based on Planck's and Einstein's contributions:
  - Energy is quantized and can only occur in discrete units called quanta
  - Electromagnetic radiation exhibits dual nature:
    - Has both characteristics of waves and particulate matter

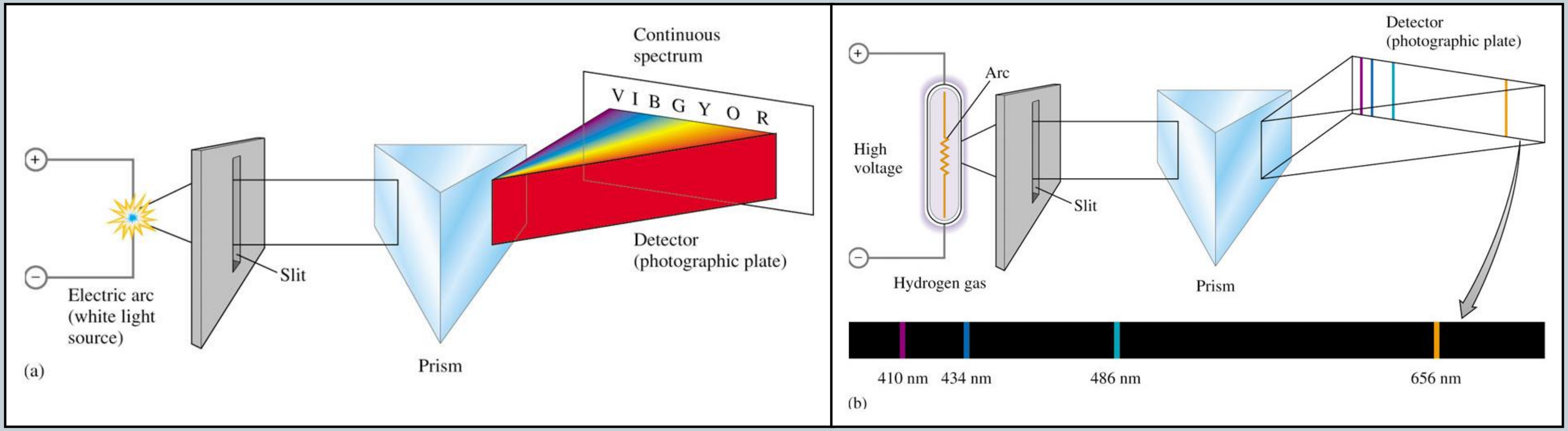
Louis de Broglie & His Contributions

- Wondered if matter displayed wavelike properties
- Used diffraction patterns to show that matter does have wave properties
- Conclusion: all matter has both particulate and wave properties

Louis de Broglie
(1892-1987)

Continuous versus Line Spectrum

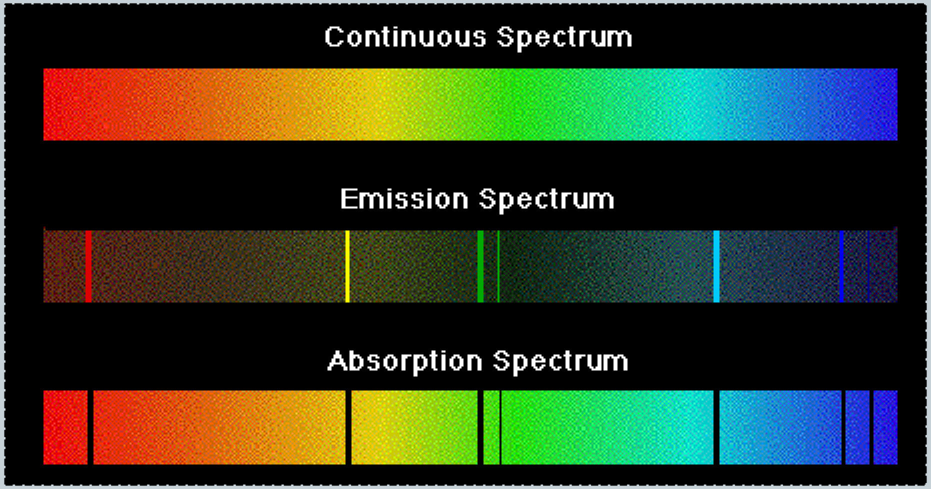
- Continuous spectrum—white light passed through a prism



- Line Spectra—only certain energies are allowed for the electrons of a hydrogen atom
  - Postulated by Max Planck

Absorption versus Emission Spectra

- Two types of line spectra:
  1) Absorption spectra—any radiation absorbed by an atom
  2) Emission spectra—certain wavelengths that excite the electrons, which then falls back to ground level, giving off light
- Absorption + Emission spectra = continuous spectra



Atomic Spectrum of Hydrogen

- Bohr's Model
  - Studied emission spectrum of Hydrogen
  - Proposed that electrons were located in energy levels around the nucleus to explain emission spectrum
  - Stated that the energy level could be calculated with the following relationship

$$E_n = \frac{-\,2.18\ x\ 10^{-18}}{n^2}\ J$$

  - When electron gains energy, it said to be in an "excited" state and then it returns to its "ground" state
    - To determine this energy gain, ΔE = energy of final state – energy of initial state

Problems with Bohr's Model

- Calculated energy levels matched with hydrogen emission spectrum
- When applied to other elements, it does not work

Atomic Spectrum of Hydrogen

- Schrodinger's Model
  - Known as the quantum mechanical model
  - Electrons are said to act like waves
  - Instead of energy levels that orbit the nucleus, it is now thought of in terms of probability
    - An orbital is the density map of where an electron could be
      - 90% probability of finding electron
    - We don't know exactly where an electron is
- Heisenberg Uncertainty Principle—we cannot know both the momentum and location of an electron

Quantum Numbers

- Used to describe the orbitals
  - Acts like coordination on a map for an electron
- Pauli exclusion principle—no two electrons can have the same four quantum numbers
- Four quantums

Four Quantum Numbers

- Principal quantum number (n)—related to the size and energy of the orbital
  - Any integers number greater than 0 (1, 2, 3, etc.)
- Angular momentum quantum number (*l*)—related to the shape of the orbital
  - Integers from 0 to n-1
- Magnetic quantum number (m_l)—related to the position of the orbital in space in relation to the other orbitals
  - Integers from $-l$ to $+l$
- Electron spin quantum number (m_s)—related to the spin of the electron
  - Only two values: +1/2 **or** -1/2

Orbital Shapes

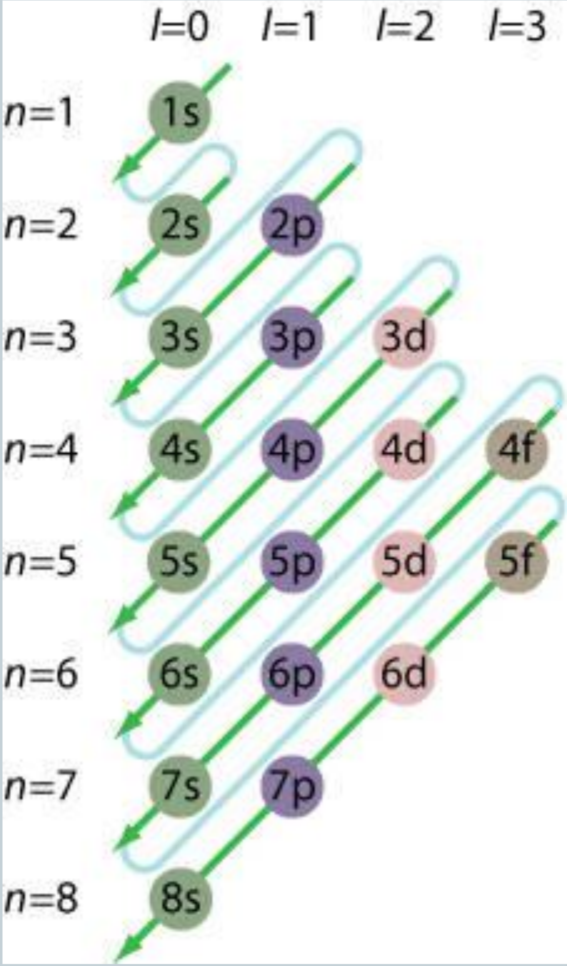
- Based on the angular momentum quantum number, each orbital is a different shape
  - Letters used to describe the different shapes

Value of *l*	0	1	2	3
Letter used	s	p	d	f
Shape	Sphere	Dumbbell	Complex	Complex

*Section 7.8 in text has p, d, & f orbital images

Electron Configuration

- Aufbau principle—scheme used to reproduce the electron configuration of the ground states of atoms by successively filling sublevels in a specific order
- Electron configuration represents all the electrons in an atom
  - Shows which energy levels and orbitals each electron is in
  - Can use the Periodic table as a tool



Orbital Diagrams

- Illustration of electron configuration now including the spin
- Hund's rule—electron have to placed in the same orbital with parallel spins before pairing electrons